Review Worksheet: Balancing Equations and Formula Mass

Balance the following equations

1. \_\_\_BaCl2 + \_\_\_K3PO4  \_\_\_Ba3(PO4)2 + \_\_\_KCl
2. \_\_\_H2O2  \_\_\_H2O + \_\_\_O2
3. \_\_\_Mg + \_\_\_ HCl  \_\_\_ MgCl2 + \_\_\_H2
4. \_\_\_NaClO3  \_\_\_NaCl + \_\_\_O2
5. \_\_\_Fe + \_\_\_H2O  \_\_\_Fe2O3 + \_\_\_H2

Calculate the molar mass of the following:

1. AlCl3

2. C3H8

3. Ca(OH)2

4. HNO3

5. PbSO4

Use the mole definition and dimensional analysis to solve the following:

1. grams to moles

530 g I2 =\_\_\_\_\_\_\_\_\_\_moles I2

1. moles to atoms

2.6 moles Cr =\_\_\_\_\_\_\_\_\_\_\_atoms Cr

1. moles to grams

1.7 moles H2S =\_\_\_\_\_\_\_\_\_\_\_g H2S

1. grams to atoms

320 g Sn =\_\_\_\_\_\_\_\_atoms Sn

1. atoms to moles

1.8 x 1023 atoms Ca =\_\_\_\_\_\_\_\_moles Ca

A recipe calls for one cup of milk and three eggs per serving. You want to make four servings for the class. How many eggs do you need? \_\_\_\_\_\_\_ Milk? \_\_\_\_\_\_

Stoichiometry \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_

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Worksheet: Mole Ratios

1. 2 KClO3  2 KCl + 3 O2

How many moles of oxygen gas are produced when 6.2 moles of KClO3 decompose?

1. 2 Cu + 1 O2  2 CuO

How many moles of cupric oxide are produced when 3.6 moles of copper are oxidized?

Balance the equation first!

1. \_\_\_Mg + \_\_\_N2  \_\_\_ Mg3N2

How many moles of magnesium are needed to produce 81 moles of Mg3N2?

1. \_\_\_ K + \_\_\_HOH  \_\_\_KOH + \_\_\_H2

How many moles of H2 are produced when 3.6 moles of K are reacted in excess water?

1. \_\_\_ C3H8 + \_\_\_O2  \_\_\_CO2 + \_\_\_H2O

How many moles of water are produced when 8.6 moles of C3H8 are combusted?

1. \_\_\_AlCl3 + \_\_\_ Na2CO3  \_\_\_Al2(CO3)3 + \_\_\_NaCl

How many moles of sodium carbonate are needed to produce 1.7 moles of Al2(CO3)3?

When in doubt, convert to moles!!

Periodic Table

Molecules/ Atoms

Moles

Chemical A

Mass

Avogadro’s #

6.02 x 1023

Periodic Table

Molecules/ Atoms

Moles

Chemical B

Mass

Avogadro’s #

6.02 x 1023

Volume of Gas @ STP

22.4 L/mol

Volume of Solid/Liquid

Density

Volume of Gas @ STP

22.4 L/mol

Volume of Solid/Liquid

Density

Worksheet: Problems Involving Mass

Use the balanced equation below to answer the questions that follow:

1. 2 Al + 1 Fe2O3  2 Fe + 1 Al2O3
2. How many grams of Al are needed to completely react with 135 g Fe2O3?
3. How many grams of Al2O3 can form when 23.6 g Al react with excess Fe2O3?
4. How many grams of Fe2O3 react with excess Al to make 475 g Fe?
5. How many grams of Fe will form when 97.6 g Fe2O3 form?
6. Phosphorus trichloride (PCl3) is made when white phosphorus (P4) reacts with chlorine gas: P4 + Cl2  PCl3
7. Balance the equation
8. How many grams of P4 are needed to produce 5.49 g PCl3? (1.24 g)
9. How many moles of oxygen are required to prepare 142 g P4O10 from elemental white phosphorus?
10. Balance the equation
11. Calculate moles of oxygen. (2.5 moles)
12. What mass of oxygen is needed for this reaction to occur? (80 g)
13. Suppose 2.17 g HgO is thermally decomposed to elemental mercury and oxygen.
14. Balance the equation
15. What mass of mercury will be produced? (2.01 g)
16. How many oxygen molecules will be produced? (3.02 x 1021 molecules)
17. How many grams of Cu2S could be produced from 9.90 g CuCl reacting with H2S gas? (Hint: This is a double replacement reaction).
18. Balance the equation
19. Find grams of Cu2S (7.96 g)
20. The action of carbon monoxide on iron(III) oxide can be represented by the equation: Fe2O3 + 3CO  2Fe + 3CO2. What would be the minimum amount of carbon monoxide used if 18.7 g of iron were produced? (14.1 g CO)

Worksheet: Theoretical Yield

* 1. How many grams of hydrogen can be produced from the reaction of 11.5 g of sodium with an excess of water? (0.505 g H2)
  2. An excess of nitrogen reacts with 2.0 g of hydrogen. How many grams of ammonia are produced? (11.2 g NH3)
  3. How many grams of CO2 will be formed when 85.6 grams of carbon are burned completely? (314 g CO2)
  4. In the decomposition of potassium chlorate, 64.2 grams of oxygen are formed. How many grams of potassium chloride are produced? (99.7 g KCl)
  5. How many grams of bromine (V) fluoride would form in the composition reaction between 384 g bromine and excess fluorine gas. (841 g)
  6. How many grams of hydrogen are produced when 5.62 grams of aluminum reacts with excess hydrochloric acid? (0.631 g H2).

Worksheet: Problems Involving Volume

Use the densities and balanced equation provided to answer the questions that follow. (density of C5H12 = 0.620 g/mL; density of C5H8 = 0.681 g/mL; density of H2 = 0.0899 g/L)

1. 1 C5H12 (l)  1 C5H8 (l) + 2 H2 (g)
2. How many milliliters of C5H8 can be made from 366 mL C5H12?
3. How many liters of H2 can form when 4.53 x 103 mL C5H8 form?
4. How many milliliters of C5H12 are needed to make 97.3 mL C5H8?
5. How many milliliters of H2 can be made from 1.98 x 103 mL C5H12?
6. What volume of hydrogen at STP can be produced from the reaction of 6.54 grams of zinc with hydrochloric acid? (single replacement rxn) (2.24 L)
7. How many grams of sodium chloride can be produced by the reaction of 112 ml of chlorine at STP with excess sodium? (composition rxn) (0.585 g)
8. An excess of hydrogen reacts with 14 grams of nitrogen. How many liters of ammonia will be produced at STP? (composition rxn) (22.4 L)
9. How many liters of oxygen are required to burn 1.00 liter of methane, CH4? (combustion rxn) (2 L)
10. How many liters of carbon dioxide will be produced by burning completely 5.00 liters of ethane, C2H6? (combustion rxn) (10 L)

Worksheet: Percentage Yield

1. A student was preparing copper metal by the reaction of 1.274 g of copper (II) sulfate with zinc metal. She isolated a yield of 0.392 g of copper. What was her theoretical and percent yield of copper? (0.5072 g Cu, 77.3% yield)

2. A student reacted 3.22 grams of sodium bicarbonate with an excess of hydrochloric acid. They produced 2.35 grams of sodium chloride. The gases carbon dioxide and water vapor were also produced. What was the student’s theoretical yield and percent yield of sodium chloride? (2.24 g, 104.9% yield)

3. A lab group reacted 0.75 grams of nails (iron) with copper(II) chloride to produce copper and iron(II) chloride. The lab group produced 0.77 grams of copper. Calculate the theoretical yield and percent yield of copper. (0.85 g, 90.6% yield)

4. A lab group reacted 0.92 grams of nails (iron) with copper (II) sulfate to produce copper and iron (II) chloride. This lab group produced 1.01 grams of copper. Calculate the theoretical yield and percent yield of copper. (1.05 g Cu; 96.2% yield)

\*\*\*5. A sample of lime, CaO weighing 69 g was prepared by heating 131 g of limestone (95% CaCO3). Carbon dioxide is also produced. What was the theoretical yield and percent yield of the reaction? (69.72 g CaO; 99% yield)

Worksheet: Limiting Reactants

A bicycle requires 1 frame and 2 wheels. A mechanic has 10 frames and 16 wheels. How many bicycles can he build? \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_

1. For the following reaction, identify the limiting reagent and the theoretical yield of phosphorous acid, H3PO3 in each scenario

PCl3 + 3H2O  H3PO3 + 3HCl

1. 3.00 mol PCl3 is mixed with 3.00 mol H2O
2. 75.0 g PCl3 is mixed with 75.0 g H2O
3. 225 g PCl3 is mixed with 125 g H2O
4. 2.00 mol PCl3 is mixed with 100 g H2O
5. If 6.57g iron reacts with 10.7 g HCl, then H2 and iron(II) chloride are produced. Determine which reactant is in excess? How much should have been put in? How much was wasted(in excess)? Determine the mass of each product. Show that the law of conservation of matter applies in this reaction!!

(HCl is in excess. 8.57 g HCl should have been put in. 2.13 g HCl was wasted(in excess). 0.2376 g H2 , 14.91 g FeCl2)

1. When 8.76 g Al reacts with 29.37 g HCl, then aluminum chloride and hydrogen are produced. Which reactant is in excess? How much should have been put in? How much was wasted(in excess)? Calculate the mass of each product. Show that the law of conservation of matter applies to this reaction.

(Al is in excess. 7.24 g should have been put in. 1.52 g of Al was wasted(in excess). 35.8 g AlCl3 and 0.8136 g H2 )

PRACTICE TEST #1

Worksheet: Everything

1. A compound was known to be either CuCl2 or CuBr2. A 5.00 g sample yielded 2.36 g of copper. What was the compound?

2. Calculate the formula of a compound, given that 55.85 g of iron combines with 32.06 g of sulfur.

3. How much iron could be obtained from 1 ton of iron ore containing 45% Fe2O3? (Put your answer in pounds of Fe)

4. How much SO2 could be obtained from burning 25 g sulfur in oxygen? (Put your answer in moles of SO2)

5. What would be the mass of the residue if 8.375 g of U(SO4)2.9H2O were heated until the water had evaporated?

6. The element M forms the chloride MCl2. This chloride contains 75% chlorine. Calculate the atomic mass of M.

7. How many tons of Fe2O3 will contain 12 tons of Fe?

8. A compound is either ZnBr2 or ZnI2. An 8.00 g sample yielded 1.64 g of zinc. What is the compound?

9. How many pounds of KCl will be formed if 50 pounds of KClO3 are decomposed by heating?

10. It was found that 10.0 g of a pure compound contains 3.65 g K, 3.33 g of Cl, and 3.02 g of O. Calculate the empirical formula of the compound. Its molecular mass is 106.55. What is its molecular formula?

Answers:

1. CuCl2 6. 23.65 g

2. FeS 7. 17.16 tons

3. 629.29 lbs. 8. ZnI2

4. 0.78 moles 9. 30.4 lbs.

5. 6.082 g 10. KClO2, KClO2

PRACTICE TEST #2

Worksheet: Everything

1. What is the key conversion factor needed to solve all stoichiometry problems?

Title: Mole Relationship in a Chemical Reaction Lab

Purpose: In this experiment, you will test the Law of Conservation of Matter by causing a reaction to occur with a given amount of reactant. You will then determine the mass of one of the products. You will then compare the experimental and theoretical mol:mol ratio between one of the reactants and one of the products. Additionally, you will determine which reactant is in excess and which is limiting. Once the excess reactant is determined, you will determine how much should have been used and how much was wasted(in excess). Then, you will compare the experimental yield of sodium chloride compared to the theoretical yield of sodium chloride produced. Finally, you will determine the percent yield of sodium chloride. This will be used to determine your accuracy grade.

Directions:

1. Clean an evaporating dish and rinse it with water from the sink. Dry it thoroughly with paper towels.
2. Obtain the mass of the evaporating dish and the watch glass to the nearest 0.01 grams.
3. With a scoopula, add about 3 grams of sodium bicarbonate (NaHCO3) to the evaporating dish and read the mass to the nearest 0.01 grams.
4. Obtain about 8 mL of 6M hydrochloric acid in a clean graduated cylinder. Record the exact measurement.
5. Slowly add the acid to the sodium bicarbonate from your graduated cylinder. Allow the drops to enter the lip of the evaporating dish so that they flow down the side gradually.
6. Continue adding all of the acid slowly. Make sure to keep the watch glass on.
7. Tilt the dish from side to side to make sure the liquid has reached the entire solid.
8. Heat the liquid in the evaporating dish over a Bunsen burner. Make sure to keep the watch glass on top of the evaporating dish. Be careful of splatter. You should be using a blue flame and waving it underneath the evaporating dish back and forth.
9. Make sure it appears to be excessively dry. Heating should take approximately 10 to 15 minutes.
10. Remove the heat from under the dish and let it cool.
11. Record the new mass to the nearest 0.01 grams.

Data:

1. Mass of empty evaporating dish + watch glass \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_

2. Mass of dish + watch glass + NaHCO3 \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_

3. Mass of NaHCO3 \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_

4. Mass of dish + watch glass + NaCl \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_

5. Mass of NaCl \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_

6. Volume of HCl \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_

Calculations/Analysis Questions:

1. Write the balanced equation that occurred in this experiment.

2. What is the theoretical mol:mol ratio between the sodium bicarbonate and the sodium chloride.

3. Calculate the experimental mass of sodium bicarbonate used and sodium chloride produced.

4. Change the mass to moles to determine the experimental mol:mol ratio. Compare this to the theoretical. How close did you come? Did you prove the Law of Conservation of Matter? Was matter created or destroyed?

5. Now take the volume of HCl and using the idea of molarity to calculate the number of moles of HCl used. Notice the theoretical mol:mol ratio between the sodium bicarbonate and HCl Compare this to the experimental mol:mol ratio between the sodium bicarbonate and HCl and determine which was in excess. Was HCl in excess? Now calculate how much HCl should have been used comparing it to the moles of sodium bicarbonate that was used. How much HCl was wasted(in excess) in ml ?

6. Now using the limiting reactant(sodium bicarbonate), calculate the theoretical yield of sodium chloride that should have been produced.

7. Using the experimental and theoretical values of sodium chloride, calculate the percent yield of sodium chloride produced.

For the “A”

8. Research sodium bicarbonate, hydrochloric acid, and sodium chloride. Find physical properties, chemical properties, and uses for each of these compounds.

Chemistry - Nail Lab

# Iron(II) or (III) – Copper (II) chloride Reaction. Iron (II) or (III) – Copper (II) sulfate reaction.

## Purpose

The purpose is to determine the ratio of copper produced to iron consumed in a single replacement reaction. Determine the type of iron(?)chloride formed in this reaction. Is it iron +2 or iron +3?

## Procedure

**Day 1**

1. Label, then mass a 150 mL beaker.

2. Put between 6.0 and 8.0 g of copper(II) chloride crystals in the beaker OR put between 3.50 and 4.00 g of copper (II) sulfate in the beaker. Check with your teacher as to which chemical we are using this year. But, you must record the exact amount you are using.

3. Add about 50 mL distilled water to the beaker. Stir to dissolve the solid with a stirring rod.

4. Mass 2 or 3 nails together to ± 0.01g.

5. Place the nails in the copper(II) chloride solution. Your teacher will show you how to do this so the reaction will continue to occur overnight. Observe the reaction; record your observations. Place the labeled beaker in the place designated by your teacher.

**Day 2**

6. Remove the nails. Rinse and scrape all the precipitate (copper metal) from the nails into your labeled 150 mL beaker. Clean the nails with steel wool and mass them again to determine the mass lost by the nails. Return the nails back to the teacher.

7. Decant solution from the 150 mL beaker. Rinse the precipitate with about 25 mL of distilled water. Let the copper settle to the bottom and then decant trying to lose as little of the solid copper as you can. Your teacher will show you how to do this. Rinse with distilled water three times making sure to decant each time but keeping as much of the copper as possible. Make sure you rinse and decant at least three times. Then place the labeled beaker in the designated area to dry overnight.

8. What are we trying to rinse away so we can get just the copper product??

### Day 3

9. Mass the beaker + dry copper. Discard the copper in the place designated by your teacher. Wash your beaker and let dry.

## Data:

|  |  |
| --- | --- |
| Mass 150 mL beaker |  |
| Mass 150 mL beaker + copper(II) chloride |  |
| Mass nails before reaction |  |
| Mass nails after reaction |  |
| Mass 150 mL beaker + dry copper |  |

## Calculations:

1. Determine the mass of copper produced and the mass of iron used during the experimental reaction.

2. Calculate the moles of iron and moles of copper involved in the reaction.

3. Determine the experimental ratio moles of iron : moles of copper

4. State the two possible reactions that could occur by writing two balanced equations. Iron reacting with the copper(II)chloride(or copper(II)sulfate). One represents iron +2 and the other represents iron +3. Determine which reaction occurred by comparing the experimental mol:mol ratio above to the theoretical mol:mol ratio from the balanced equation.

5. Determine the amount of copper(II) chloride crystals(or copper(II)sulfate) that should have reacted with the amount of iron that actually reacted. Perform a mass-mass problem starting with mass of iron changing to mass of copper(II) chloride crystals(or copper(II) sulfate depending on which chemical used). Make sure you use the correct theoretical mol:mol ratio for this conversion.

6. Determine how much copper(II) chloride(copper(II)sulfate) crystals were wasted by subtraction. Initial copper crystals used – copper crystals that actually reacted = copper crystals wasted.

7. Determine the theoretical amount of copper that should have been made(Theoretical yield). Again, perform another mass-mass problem starting with grams of iron and calculating the grams of copper.

8. Determine what was the percent yield of copper. (Experimental yield of copper / theoretical yield of copper) x 100

**Conclusion:**

1. Why did the reaction stop? Which reactant was used up? How do you know?

2. Describe what was happening to the atoms of iron and copper during the reaction. What is this type of reaction called?

3. What would happen to the ratio of iron to copper if you had placed more nails in the beaker?

What would happen to the ratio of iron to copper if you let the reaction go for less time?

4. What is the accepted ratio of iron atoms to copper atoms in this reaction.? Account for differences between your experimental value and the accepted value.

For the “A”

Research iron, copper, copper(II) chloride (or copper(II) sulfate if that was used). State physical properties, chemical properties and uses for each of the substances.